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Chemistry
Teach Yourself Series
Topic 2: Titrations

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Contents

Acid/Base Titrations

What is a titration?

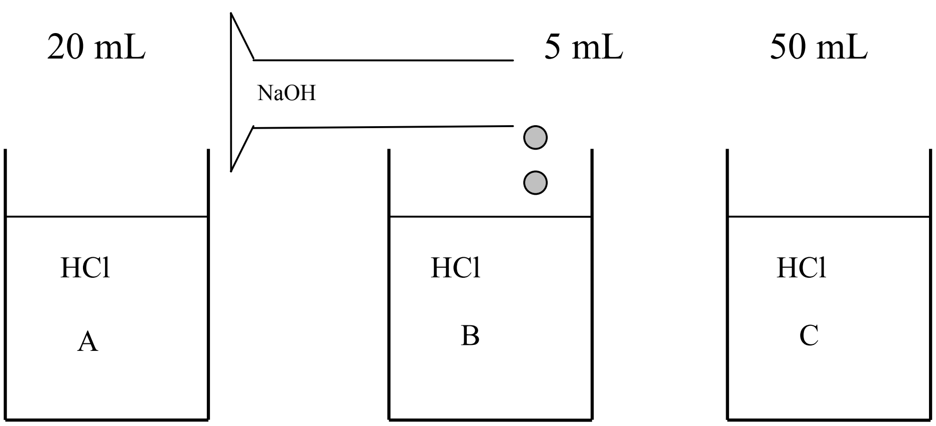
As it appears in Unit 2

In an acid/base titration, the concentration of a solution is determined by reacting the solution with another solution of known concentration.

The concentration of an acid is determined by its reaction with a base solution of known concentration.

The concentration of a base is determined by its reaction with an acid solution of known concentration.

The Principle of a Titration



Three beakers each contain 40 mL of hydrochloric acid, but the concentration of the acid in each beaker is unknown. A few drops of indicator are added to each beaker and 0.100 M sodium hydroxide is added to each beaker. The volume of sodium hydroxide required for a colour change is shown for each beaker.

Review Questions

Which beaker contains the strongest acid?
What is the ratio of the concentrations present?
What is the actual concentration of each beaker?

Beaker C contains the acid of highest concentration, then beaker A, then B.
Beaker A required 20 mL to neutralize 40 mL of acid. Therefore the base was twice as concentrated as the acid. The acid must be 0.0500 M, since the base is 0.1 M.

Therefore, the acid in Beaker B is ¼ the concentration of Beaker A as 5 mL is a ¼ of 20 mL. Acid concentration in Beaker B is 0.0125 M.

Beaker C contains acid that is 2 and a ½ times stronger than Beaker A (20:50), or 10 times stronger than Beaker B (5:50). Concentration in C = 0.125 M

In summary, the more base required, the stronger the concentration of the acid.
A titration using a burette is just a more accurate way of obtaining the volume of solution added.

Titration Setup

The volume in the burette must be read from the **bottom of the meniscus**.

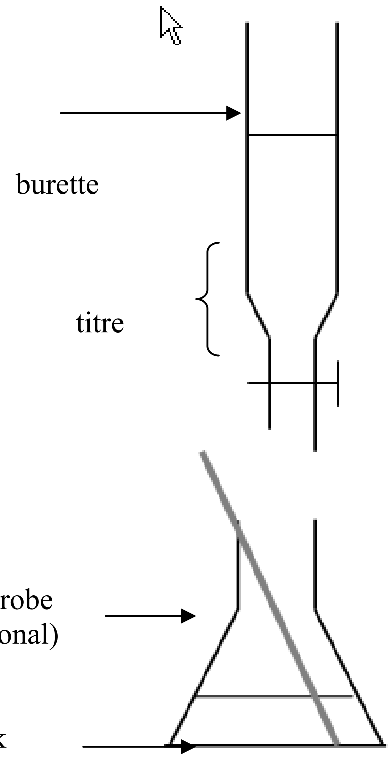
The **titre** is the volume added from the burette before the colour change.

When the colour changes, you have reached the **endpoint** and you should stop adding from burette.

A pH probe can be used instead of indicator

A standard volume of one solution (**aliquot**) is added to a flask

A few drops of **indicator** are added to the flask or a pH probe is used.



The greater the titre, the stronger the concentration of the solution in the flask.
The volume of the titre indicates the strength of the solution in the flask.

Example 1

A 25.00 mL aliquot of 0.100 M sodium hydroxide is added to a flask. It is titrated against a solution of hydrochloric acid of unknown concentration. The average titre is 16.00 mL. Calculate the concentration of the acid.

HCl(aq)	+	NaOH(aq)	→	NaCl(aq)	+	$H_2O(l)$
c=?		c=0.100				
V= 16mL		V=25 mL				
= 0.016 L		=0.025 mL				

> **Note the procedure**:
> Write a balanced equation.
> Place the data provided under the chemical that it refers to.
> Calculate the number of mole of the solution that you have **c** and **V** for.

$n(NaOH) = c \times V = 0.1 \times 0.025 = 0.0025$ mol
$n(HCl) = n(NaOH) = 0.0025$ mol

$$c(HCl) = \frac{n}{V} = \frac{0.0025}{0.016}$$

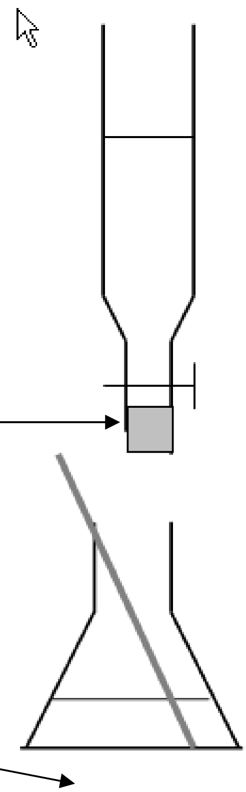
$$= 0.156 \text{ M}$$

Rinsings

Correct rinsing of each piece of equipment leads to good results.

Rinsing the burette with water would leave water in the burette.
When the solution is added to the burette, rinse the burette with the same solution as is going into it.

Rinsing the flask with water will, however, not be a problem.
The solution is added to this flask using a pipette.
The addition of water will not change the number of mole in the flask.



Choice of indicator

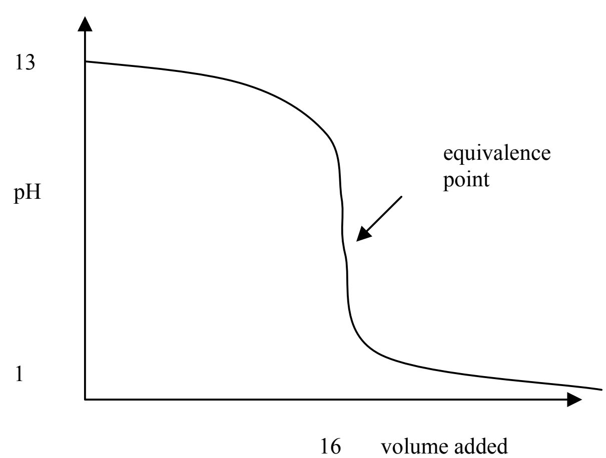
As it appears in Unit 3

NaOH in the flask is a base. The pH will be about 13. As the acid is added, the pH drops. The pH probe records the change.

As the base is close to being neutralized, the pH drops very quickly.

If too much acid is added, the pH drops to below 1.

The indicator needs to change colour at the pH of the **equivalence point**, in this case around 7.
The chemistry data book contains a list of indicator pH transition ranges.



Review Questions

1. The concentration of a sample of nitric acid is to be determined by titration against 0.100 M NaOH. 20.00 mL aliquots of NaOH are used and the average titre is 12.5 mL. Calculate the concentration of the nitric acid solution.

Balanced equation: HNO_3(aq) +
V= c= ,V=

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2. The concentration of a sample of hydrochloric acid is to be determined by titration against 0.100 M Na_2CO_3. 25.00 mL aliquots of Na_2CO_3 are used and the average titre is 17.40 mL. Calculate the concentration of the hydrochloric acid solution.

Balanced equation:
(note that this is not a 1:1 reaction)

n(HCl) = ½ n(Na_2CO_3)

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3. For Question 2, what should the
- burette be rinsed with

- flasks be rinsed with?

More advanced titrations

As it appears in Unit 3

Frequently the solution to be analysed has to be diluted before the titration.

Example 1
10.0 mL of ammonia, NH_3, solution is added to a volumetric flask and diluted to 250 mL.

The diluted ammonia is then titrated against 0.100 M hydrochloric acid solution. 20.0 mL aliquots of the ammonia solution are used and the average titre of hydrochloric acid is 24.6 mL.
Calculate the concentration of the diluted ammonia.
Calculate the concentration of the original ammonia solution.

Note: Dilution factor is $\frac{\mathbf{250}}{\mathbf{10}}$

Solution

$HCl(aq) + NH_3(aq) \rightarrow NH_4^+(aq) + Cl^-(aq)$

c = 0.1   V = 0.02 L
V =0.0246 L   c= ?

$n(\text{HCl}) = cxV = 0.1x0.0246 = 0.00246mol$

$n(\text{NH}_3) = n(\text{HCl}) = 0.00246mol$

$c(\text{NH}_3) = \dfrac{0.00246}{0.02} = 0.123M$ this is the concentration of NH_3 in flask; diluted ammonia.

conc. of original = conc. diluted x dilution factor $= 0.123x\dfrac{250}{10} = 3.08M$

Example 2
5.00 g of Draino is added to a 100 mL volumetric flask and made up to the mark with distilled water. The Draino contains NaOH as the active ingredient.
25.00 mL aliquots of the Draino solution are titrated against 0.220 M hydrochloric acid. The average titre is 12.8 mL.

Calculate the %(m/m) of sodium hydroxide in Draino

Solution

$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$

c = 0.22   V = 0.025 L
V =0.0128 L   c= ?

$n(\text{HCl}) = cxV = 0.22x0.0128 = 0.00282mol$

$n(\text{NaOH}) = n(\text{HCl}) = 0.00282mol$

$$c = \frac{0.00282}{0.025} = 0.113M$$

This time the question asks for %(m/m), therefore the mass of sodium hydroxide is required. Volume of the flask is 100 mL

n(NaOH) in 100 mL flask = c x V = $0.113x0.1 = 0.0113mol$

m(NaOH) = $nxM = 0.0113x40 = 0.452g$

$$\%(m/m)(NaOH) = \frac{\text{mass of sodium hydroxide}}{\text{mass of Draino}} \times \frac{100}{1} = \frac{0.452}{5}x100 = 9.04\%$$

Review Questions

4. Car batteries use sulfuric acid as an electrolyte. A 10 mL sample of sulfuric acid is diluted to 250 mL in a volumetric flask. 20 mL aliquots of the acid are titrated against 0.500 M sodium hydroxide solution, NaOH. The average titre is 12.8 mL. What is the concentration of the sulfuric acid?

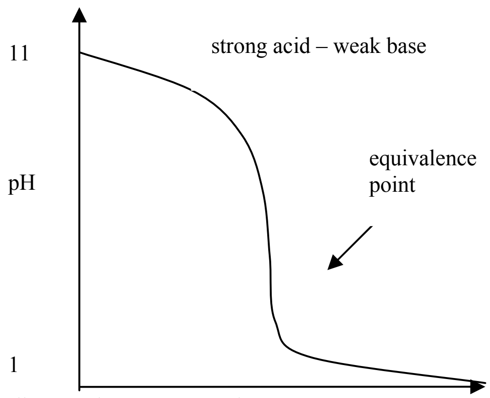
5. An impure sample of baking powder contains some sodium carbonate, Na_2CO_3. The baking powder is reacted with hydrochloric acid to determine its purity. A 1.50 g sample of baking powder is dissolved in water and added to a 250 mL flask and made up to the mark with distilled water. 20 mL aliquots of the this solution are titrated against 0.144 M HCl. The average titre is 14.3 mL.
Determine the % sodium carbonate by mass in the baking powder.

Indicator choice

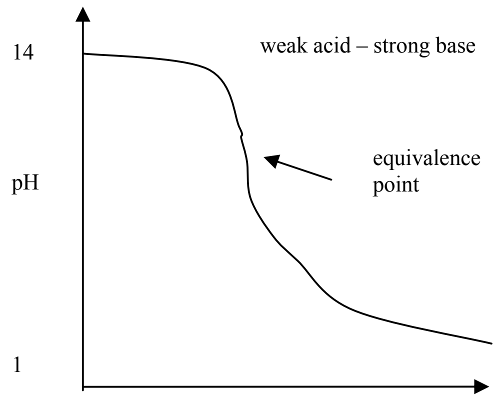
Three possibilities

1. **1.** Strong acid vs Strong base (shown above)
2. **2.** Strong acid vs Weak base
3. **3.** Weak acid vs Strong base

(N.B. Weak acids should not be titrated against weak bases)
The choice of indicator should be different for the three scenarios.



indicator changes around 8 – 9

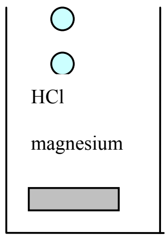


indicator changes around 3 – 4

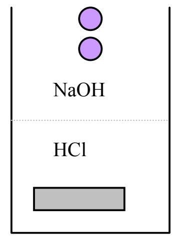
Back titrations

As it appears in Unit 3

Aim: To analyse an impure piece of magnesium.



Method 1: Add HCl drop by drop until all the magnesium has reacted. Works fine but might take hours.



Method 2: Add too much HCl. Leave to react. Then react left over HCl with NaOH. (Time issue solved)

Method 2 is a **back titration**. Too much acid is added on purpose and a second reaction is performed to mop up the excess acid. This process is easier than the direct method.

Procedure
Step 1: Excess chemical added to a typical analysis reaction.
Step 2: Work out how much in excess the chemical was.
Step 3: Determine the amount of reactant actually present in step 1.

Example
Determination of steel content in steel wool.
A 5.0 g sample of rusty steel wool is dissolved in excess sulfuric acid; 100 mL of 1.0 M sulfuric acid to be exact. (This is step one – an excess of acid is used. Step two will involve determining how much excess acid was present)
To determine the amount of sulfuric acid remaining, a titration with 0.50 M sodium hydroxide is conducted. The volume of NaOH required is 52.4 mL.
Determine the purity of the steel wool.

initial amount of sulfuric acid (excess)

Step 1: $Fe(s) + H_2SO_4(aq) \rightarrow FeSO_4(aq) + H_2(g)$
c = 1.0
V=0.10 L

Step 2: $H_2SO_4(aq) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$
c = 0.50 M
V = 0.0524 L

this base was required to mop up excess H_2SO_4

Most back titrations can be set out as a template like this.
$n(H_2SO_4$ initial$)$ = c x V = 1.0 x 0.10 = 0.10*mol*

n(NaOH) = c x V = 0.50 x 0.0524 = 0.0262*mol*
$n(H_2SO_4$ not reacting in step 1$)$ = ½ n(NaOH) = ½ x 0.0262 = 0.0131*mol*

$n(H_2SO_4$ actually reacting in step 1$)$ = $n(H_2SO_4$ initial$)$ - $n(H_2SO_4$ not reacting step 1$)$
= 0.10 – 0.0131 = 0.0869*mol*

$n(Fe)$ = $n(H_2SO_4$ actually reacting in step 1$)$ = 0.0869mol mass(Fe) = nxM = 0.0869 x 55.8 = 4.84 g

$$\%Fe\ (m/m) = \frac{4.84}{5} x \frac{100}{1} = 96.8\%$$

Review Question

6. Sand usually contains some eroded shells. The main chemical in shells is calcium carbonate, $CaCO_3$. To determine the calcium carbonate content of a sample of sand, 100 mL of 1.0 M HCl is added to a 5.0 g sample of sand.
After the reaction is completed, 0.40 M sodium hydroxide solution is used to find how much HCl did not react with the sand. 74.8 mL of NaOH is required.
Use the grid below to help you determine the $CaCO_3$ content of the sand.

Step 1 Eqn: *Write eqn here*

Write in the concentration and volume of the HCl

Step 2 Eqn: *Write eqn here for the mopping up reaction*

Write in the concentration and volume of the NaOH

Calculate initial HCl

Calculate n(NaOH)
Etc

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Solutions to questions

1. Balanced equation: $HNO_3(aq)$ + $NaOH(aq)$ → $NaNO_3(aq)$ + $H_2O(l)$)
V= 0.0125 c= 0.1,V= 0.02

$n(\text{NaOH}) = cxV = 0.1x0.02 = 0.002mol$

$n(HNO_3) = 0.002mol$

$$c = \frac{n}{V} = \frac{0.002}{0.0125} = 0.160M$$

2. Balanced equation: $2HCl(aq)$ + $Na_2CO_3(aq)$ → $2NaCl(aq)$ + $H_2CO_3(aq)$
V=0.0174 c=0.1,V=0.025

$n(Na_2CO_3) = cxV = 0.1x0.025 = 0.0025mol$

$n(\text{HCl}) = 2\text{x}n(Na_2CO_3) = 0.0050mol$

$$c(\text{HCl})= \frac{n}{V} = \frac{0.005}{0.0174} = 0.287M$$

3. - burette be rinsed with – solution in it - HCl

- flasks be rinsed with – distilled water

4. $H_2SO_4(aq)$ + $2NaOH(aq)$ → $Na_2SO_4(aq)$ + $2H_2O(l)$
c = ? V=0.0128 L
V = 0.020 c= 0.500 M

$n(\text{NaOH}) = cxV = 0.5x0.0128 = 0.0064mol$
$n(H_2SO_4) = ½\, n(\text{NaOH}) = 0.0032mol$

$$c(H_2SO_4) = \frac{n}{V} = \frac{0.0032}{0.020} = 0.16mol$$

$$c(\text{undiluted } H_2SO_4) = 0.16x\frac{250}{10} = 4.00M$$

5. $2HCl(aq) + Na_2CO_3(aq) \rightarrow 2NaCl(aq) + H_2CO_3(aq)$

c= 0.144 M   c=?

V=0.0143 L   V = 0.020 L

n(HCl) = $cxV = 0.144x0.0143 = 0.00206M$

n(Na_2CO_3) = ½ n(HCl) = ½ x 0.00206 = 0.00103*mol*

n(Na_2CO_3 in the 250 mL flask) = $nx\frac{250}{20} = 0.00103x\frac{250}{20} = 0.0129mol$

mass(Na_2CO_3) = $nxM = 0.0129x106 = 1.36g$

% (Na_2CO_3) = $\frac{1.36}{1.5}x100 = 90.9\%(m/m)$

6. Step 1: $2HCl(aq) + CaCO_3(s) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$

c = 1.0

V = 0.10

Step 2: $HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$

c = 0.40

V = 0.0448 L

n(HCl initial) = cxV= 1.0 x 0.10 = 0.10 *mol*

n(NaOH) = cxV = 0.40 x 0.0748 = 0.0299 *mol*

n(HCl in excess) = n(NaOH) = 0.0299 *mol*

n(HCl reacting with sand) = 0.10 – 0.0299 = 0.0701 *mol*

n($CaCO_3$) = ½ n(HCl) = ½ x 0.0701 = 0.035 *g*

mass($CaCO_3$) = nxM = 0.035x100 = 3.5 *g*

% $CaCO_3$ by mass = $\frac{3.5x100}{5x1} = 70\%$